

TOPIC E –

ENVIRO CHEMISTRY

PART 12 – FURTHER

WATER AND SOIL HL

IB Chemistry

Topic E – Enviro

Hodder Ed - Talbot



E12 Water and soil - 3 hours

- E.12.1 Solve problems relating to the removal of heavy-metal ions, phosphates and nitrates from water by chemical precipitation. (3)
- E.12.2 State what is meant by the term cation-exchange capacity (CEC) and outline its importance. (2)
- E.12.3 Discuss the effects of soil pH on cation-exchange capacity and availability of nutrients. (3)
- E.12.4 Describe the chemical functions of soil organic matter (SOM). (2)



E12.1 – Chemical Precipitation

- E.12.1 Solve problems relating to the removal of heavy-metal ions, phosphates and nitrates from water by chemical precipitation. (3)
- *Given the equilibrium formed by a metal M and a non-metal X : $MX(s) \rightleftharpoons M^+(aq) X^-(aq)$.*
- *The K_{eq} for this system is given by $K_{sp} = [M^+][X^-]$, and is called the solubility product constant.*
- *Students should be able to solve problems associated with this type of equilibrium, including the common ion effect.*



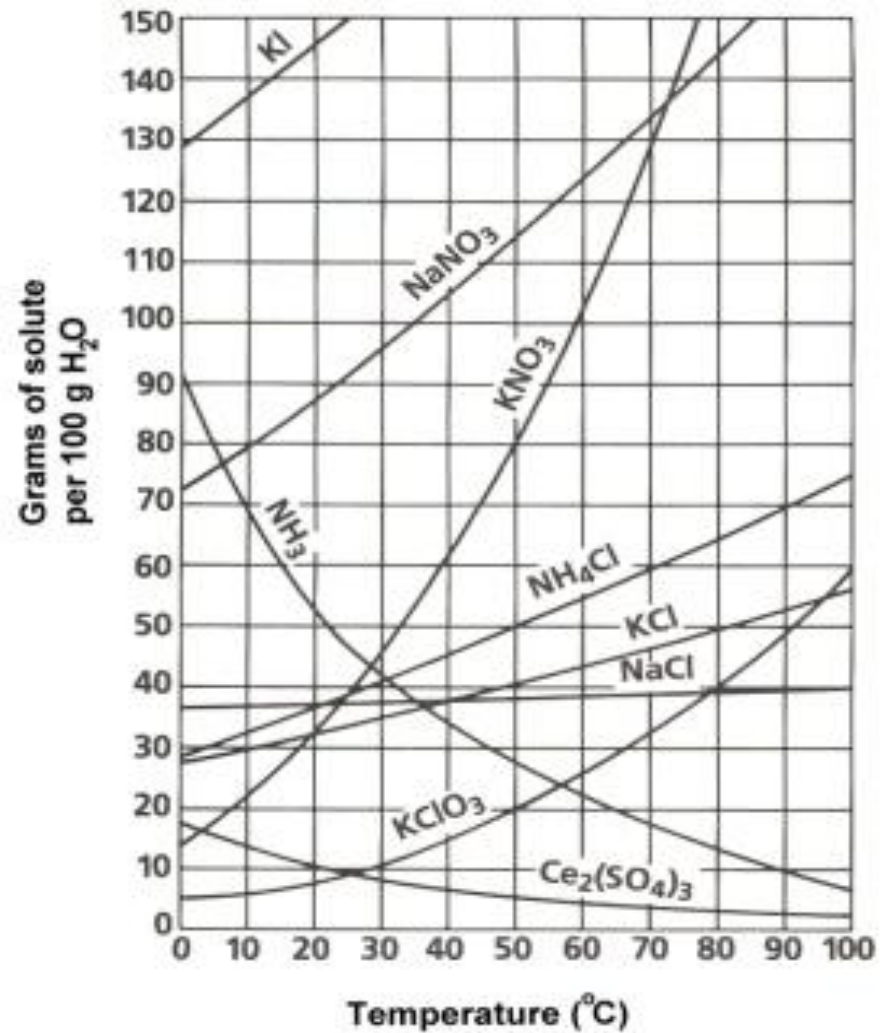
E12.1 – Precipitation by Temp

- Remember, from Part 06, the limited solubility of transition metal hydroxides allows us to remove the transition metals from waste water by chemical precipitation. (the metals can inhibit enzymes)
- Unwanted soluble materials precipitate out of solution when they pass the saturation point at a given temperature.
 - By addition of other substances, this point of solubility can be lowered even further
 - A form of the equilibrium constant (K_c) called the solubility product (K_{sp}) is then used.



E12.1 – Solubility Curves

- Increased T causes the saturation point to increase for solids and decrease for gases
- Additional substances added to a solution can lower the solubility of unwanted compounds in water



E12.1 – Producing K_{sp}

- Consider partially soluble salt, MX
 - Will dissolve to a limited degree, in equilibrium
 -
- The equilibrium expression is:
 - $$\frac{[\text{ } (\text{ })] (\text{ })}{[\text{ } (\text{ })]}$$
- Heterogeneous equilibrium, $[\text{MX}](s)$ is constant at a given temperature, rearranging gives the **solubility product constant, K_{sp}**



E12.1 – Solubility of Metals

- For a complex transition metal ion such as M_3X_2 :
 - $M_3X_2(s) \rightleftharpoons yM^{2+}(aq) + zX^{3-}(aq)$
 -
- For aluminum hydroxide, $Al(OH)_3$:
 - $Al(OH)_3(s) \rightleftharpoons Al^{3+}(aq) + 3OH^-(aq)$
 -
- This is a constant at a given temperature, **changing temperature will change K_{sp}** and the amount of substance that will dissolve in given conditions



E12.1 – Calculating K_{sp} from solubility

- If the solubility (—) is known, K_{sp} can be calculated
- Solubility of lead (II) bromide at 298K is 6.15—
 - $K_{sp} = [Pb^{2+}(aq)][Br^{-}(aq)]^2$
- You must calculate the molarity (—) of $PbBr_2$
 - $PbBr_2 = \text{—}$
 - $[Pb^{2+}(aq)] = \text{—}$ $[Br^{-}(aq)]^2 = \text{—}$
 - $K_{sp} = 0.0214 \times (0.428)^2 = 3.9 \times 10^{-5}$



E12.1 – Calculating Solubility from K_{sp}

- $K_{sp} [\text{Al}(\text{OH})_3] = 1.0 \times 10^{-32}$
- - Dissociation of 1mol $\text{Al}(\text{OH})_3$ produces 1mol of $\text{Al}^{3+} = x$ ions and 3mol $\text{OH}^- = 3x$ ions
 -
- We can now substitute the value for K_{sp}
 -
 -
 - — (solubility of $\text{Al}(\text{OH})_3$)
 - The solubility in grams per dm^3



E12.1 – Common Ion Effect

- The **common ion effect** results when one of its constituent ions is already present in solution and therefore the salt is less soluble
- Example AlCl_3 in solution, then add NaCl
 - $\text{AgCl(s)} \rightleftharpoons \text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
 - With the addition of another salt, NaCl , which also contains $\text{Cl}^-(\text{aq})$ ions and has its own equilibrium ($\text{NaCl(s)} \rightleftharpoons \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$)
 - Le Chatelier's principal (Topic 07) predicts that the equilibrium of AlCl_3 will shift back left to a lower $[\text{Cl}^-]$ ions than before.



E12.1 – AgCl in Water

- Consider the following example:
- For AgCl $K_{sp}=2.0 \times 10^{-10}$ at 298K, dissolved in H₂O
 - $\text{AgCl(s)} \rightleftharpoons \text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
 - $K_{sp} = [\text{Ag}^+(\text{aq})][\text{Cl}^-(\text{aq})]$ (where they are equal, and x)
 - $K_{sp} = x^2 = 2.0 \times 10^{-10}$
 - $x = (2.0 \times 10^{-10})^{1/2}$
 - $x = 1.4 \times 10^{-5} \text{ mol/dm}^3$ (molarity of AgCl in water)



E12.1 – AgCl in NaCl solution

- If the same salt is dissolved in an NaCl solution instead of H₂O
- $\text{AgCl(s)} \rightleftharpoons \text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
- $\text{NaCl(s)} \rightleftharpoons \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$ (solution is 0.1M)
- Since there are now two sources of Cl⁻ ions, we will give the molarity of Ag⁺ ions (y), so Cl⁻ ions is (0.1 + y).
- $K_{\text{sp}} = [\text{Ag}^+(\text{aq})][\text{Cl}^-(\text{aq})] = y + (0.1 + y)$
 - The contribution of Cl⁻ ions from NaCl is much greater than that from AgCl since AgCl is only slightly soluble, where NaCl is very soluble
- $y \ll 0.1$ hence $(0.1 + y) = 0.1$
- $K_{\text{sp}} = 0.1 \times y = 2.0 \times 10^{-10}$
- $y = 2.0 \times 10^{-9} \text{ mol/dm}^3$



The presence of NaCl, lowered the solubility of Cl⁻ ions from AgCl in solution considerably

E12.1 - Precipitation

- Remember, precipitation reactions are used to remove heavy metal pollutants from solutions
 - $\text{Fe}^{3+}(\text{aq}) + \text{PO}_4^{3-} \rightarrow \text{FePO}_4(\text{s})$
 - $\text{Al}^{3+}(\text{aq}) + \text{PO}_4^{3-} \rightarrow \text{AlPO}_4(\text{s})$
- The precipitation can be predicted if the ionic product is greater than the solubility product for a given ion
- Please see the worked example in your Talbot IB Chemistry Book, also check out the explanation for EDTA simply for Example:



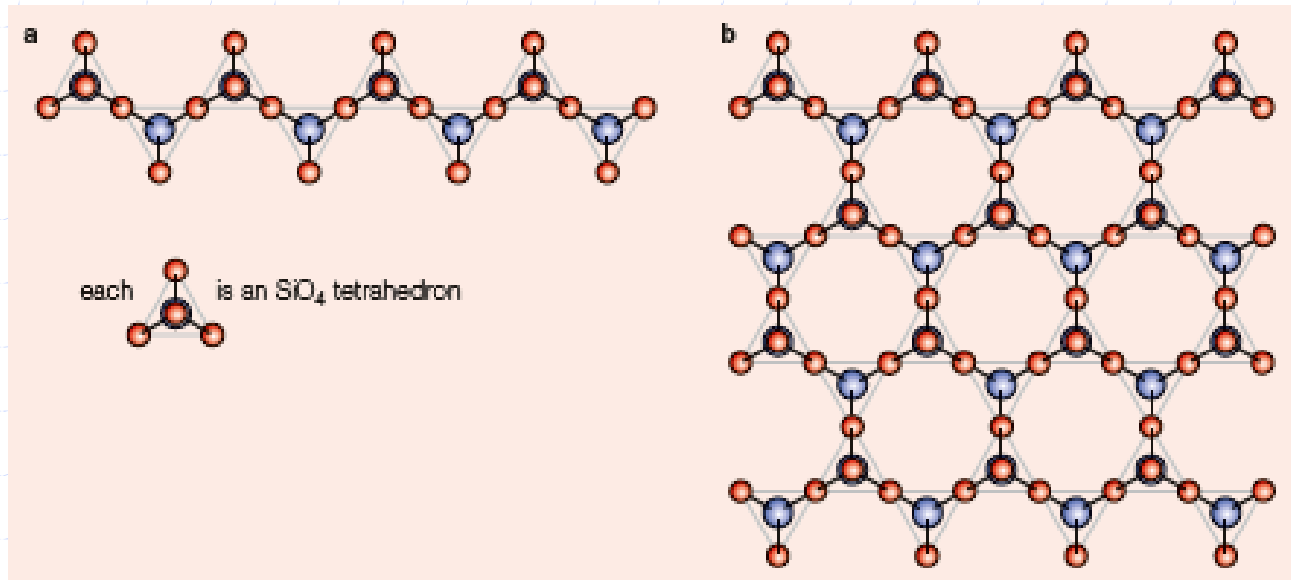
E12.2 – Cation-exchange

- E.12.2 State what is meant by the term cation-exchange capacity (CEC) and outline its importance. (2)
- *The amount of exchangeable cations in a clay is called cation-exchange capacity. Include equations as appropriate.*



E12.2 – Clay Soil

- Clay soils are able to retain water and nutrient ions which might otherwise be leached out of the soils by rain water.
 - The minerals in clay carry a negative charge that is balanced positive metal ions



E12.2 – Cation-exchange capacity (CEC)

- The cations bound to the clay structure are not permanent as they can be exchanged
- **Cation-exchange capacity** (CEC) is defined as the amount of the single-positive cations that can be exchanged with the soil solution, per Kg of clay
 - If the clay contains Mg^{2+} or Al^{3+} , these ions must be divided by their charge
 - $\text{Na}^+ + 1/2 \text{Mg}^{2+} + 1/3 \text{Al}^{3+}$
 - CEC values range from 0.03 —for kaolinite up to 1.5— for vermiculite clay.



Most soils contain small fraction of clay, so CEC's for soils range from 0.02 – 0.60 —

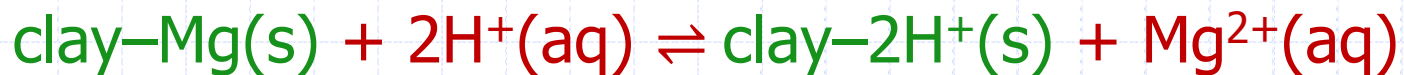
E12.3 – Effect of pH on CEC

- E.12.3 Discuss the effects of soil pH on cation-exchange capacity and availability of nutrients. (3)
- *Examples of nutrients include Ca, Mg, Fe, Al, P, N, S, Cu and Zn. Use equations as appropriate.*
- When low pH solutions (acids with plenty of H^+) are added to soils, the equilibrium changes and the H^+ take place in the soil, releasing nutrients retained by the soil into the ground water to be washed away.



E12.2 – CEC Example

- A typical exchange process:
$$\text{clay-Mg(s)} + 2\text{NH}_4^+(\text{aq}) \rightleftharpoons \text{clay-2NH}_4^+(\text{s}) + \text{Mg}^{2+}(\text{aq})$$
- This equilibrium can shift in response to changing concentrations in the soil of Mg^{2+} and NH_4^+ ions as a result of chemical additions to the soil such as fertilizers or the use of ions in plant growth
- Without this ability to exchange and retain ions, soil would not be able to supply nutrients to its surroundings



E12.3 – low pH effect on Anions in Soil

- Low pH also affects the availability of anion nutrients such as phosphate, PO_4^{3-} , and nitrate, NO_3^- . Now, the low pH increases the ions (since they are negative)



- In addition, as we talked about, low pH enables the reduction of NO_3^- ions to NH_4^+ ions



- Plants take up nitrogen in the form of NO_3^- ions, so this process lowers the availability of nitrogen for plants to take up



E12.4 – Chemical functions of SOM

- E.12.4 Describe the chemical functions of soil organic matter (SOM). (2) *Include the following:*
 - *Contributes to cation-exchange capacity*
 - *Enhances the ability of soil to buffer changes in pH*
 - *Binds to organic and inorganic compounds in soil*
 - *Reduces the negative environmental effects of pesticides, heavy metals and other pollutants by binding contaminants*
 - *Forms stable complexes with cations*



E12.4 – Props of SOM in Soil

- The action of anaerobic bacteria on organic matter in the soil leads to the formation of phenols, carboxylic acids, proteins and peptides.
- These substances lend the soil certain chemical properties.
 - Binding metal cations in soil
 - Increasing cation-exchange capacity of soil
 - Binding organic pollutants
 - Increasing the buffering capacity of the soil (resistance to pH changes)



E12.4 – Read it up!

- For further information on the functions of soil adopted from soil organic matter, please read page 918 of Talbots book.

